Name: ______Key______

Chemistry 129.03 Spring 2011

General Chemistry

Examination #3:

Equations are provided.

You may use a calculator.

Show all your work!

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Bonus: ______/2

Total: ______/100

1. (a) Classify the following salts as basic, acidic or neutral. (6 pts.)

(i) KCl	neutral
(ii) NH_4NO_3	acidic
(iii) $NaHSO_3$	acidic
(iv) $Ba(C_2H_3O_2)_2$	basic
(v) $Fe(ClO_4)_3$	acidic
(vi) Na_2CO_3	basic

(b) Rank the following acids in order of increasing strength. Explain. (6 pts)

$HClO_2$, $HClO_4$, $HClO$, $HClO_3$

$$HClO < HClO_2 < HClO_3 < HClO_4$$

For oxyacids with the same central atom, acid strength increases as the number of oxygen atoms attached to the central atom increases.

2. (6 pts.) A 1.0 L buffer solution that is 0.125 M $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$) and 0.100 M in $NaC_2H_3O_2$ has a pH of 4.65. What is the pH of the buffer solution after 0.030 mol of HCl have been added?

$$NaC_2H_3O_2 + HCl \longrightarrow HC_2H_3O_2 + NaCl$$

	$NaC_2H_3O_2$	HCl	$HC_2H_3O_2$	$NaCl$
BA	0.100 mol	—	0.125 mol	
A	—	0.030 mol	—	
AA	0.070 mol		0.155 mol	

$$pH = pK_a + \log\frac{[base]}{[acid]}$$

$$pH = -\log(1.8\times10^{-5}) + \log\frac{(0.070)}{(0.155)} = \underline{4.39}$$

3. (16 pts) Consider the titration of 15.00mL of 0.200 M ammonia (NH_3), $K_b = 1.8 \times 10^{-5}$, with 0.100 M HCl:

$$NH_{3\,(aq)} + HCl_{(aq)} \rightarrow NH_4Cl_{(aq)}$$

Determine:

(a) the [OH^-] and the initial pH of the ammonia solution (6 pts.)

$[OH^-] = \sqrt{[NH_3] \times K_b}$

$[OH^-] = \sqrt{(0.200)(1.8\times10^{-5})} = \underline{1.9\times10^{-3} M}$

$pOH = -\log[OH^-] = -\log(1.9\times10^{-3}) = 2.72$

$pH = 14.00 - pH = 14.00 - 2.72 = \underline{11.28}$

(b) the [H_3O^+] and pH after the addition of 30.00mL of HCl. What point in the titration is this? Does the pH value make sense? Why? (10 pts.)

Equivalence Point

$[H_3O^+] = \sqrt{[NH_4Cl] \times Ka}$

$[H_3O^+] = \sqrt{(5.6\times10^{-10})(0.0667)} = \underline{6.1\times10^{-6} M}$

$pH = -\log[H_3O^+] = -\log(6.1\times10^{-6}) = \underline{5.21}$

$Ka = \frac{Kw}{Kb} = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$

$[NH_4Cl] = \frac{3.00\ mmol}{45.00\ mL} = 0.0667 M$

	NH_3	+ HCl →	NH_4Cl
BA	3.00 mmol	—	—
A	—	3.00 mmol	—
AA	0	0	3.00 mmol

*pH is below 7 because at the equivalence point there's a NH_4Cl solution which is an acidic salt.

4. An ammonia (NH_3) solution has a pH of 11.13. Which of the following substances will decrease the pH of the solution upon addition? (4 pts)

(a) $Ca(NO_3)_2$ (b) NH_4Cl (c) KI (d) NaOH

5. (12 pts) A gas is confined to a cylinder under constant atmospheric pressure. When 378 J of heat are absorbed by the gas it expands and does 56 J of work on the surroundings.

(a) What is the change in the internal energy of the system? What is the change in the internal energy of the surroundings? (6 pts)

$q = +378\ J$ $w = -56\ J$

$\Delta E = q + w = 378J + (-56J) = \underline{322\ J}$

↑ system

$\Delta E_{surr} = \underline{-322\ J}$

(b) What is the change in the enthalpy? Why? (3 pts)

$\Delta H = q_p$ @ constant pressure

$\Delta H = \underline{+378\ J}$

(c) Is the process endothermic or exothermic? Why? (3 pts)

Endothermic Process because heat is absorbed.

6. (a) (3 pts) The same quantity of heat is added to 10.00 g of gold, 10.00 g of iron, and 10.00 g of copper, all initially at 25 °C. The specific heats of these metals are:

$$Au = 0.125 \frac{J}{g\ ^{\circ}C} \qquad Fe = 0.460 \frac{J}{g\ ^{\circ}C} \qquad Cu = 0.397 \frac{J}{g\ ^{\circ}C}$$ (3 pts)

Which piece of metal has the highest final temperature? Why?

Au because it has the smallest specific heat capacity.

(b) (5 pts) Calculate the final temperature of 10.00 g of this metal after 100.0 J of heat have been added.

$$q = m \times C_s \times \Delta T$$

$$100.0\,J = (10.0g)(0.125 \frac{J}{g \cdot ^{\circ}C})(\Delta T)$$

$$\Delta T = 80.0^{\circ}C = T_f - T_i$$

$$T_f = 80.0^{\circ}C + 25^{\circ}C = \underline{105^{\circ}C}$$

7. (9 pts) Calculate ΔH_{Rxn} for the reaction :

$$2\ CH_{4\,(g)} + 3\ O_{2\,(g)} \rightarrow 2\ CO_{(g)} + 4\ H_2O_{(g)} \qquad \Delta H = ?$$

given the following set of reactions and their respective enthalpy changes:

$$CO_{2\,(g)} + 2\ H_2O_{(g)} \rightarrow CH_{4\,(g)} + 2\ O_{2\,(g)} \qquad \Delta H = +802\ kJ$$
$$CO_{(g)} + \tfrac{1}{2}\ O_{2\,(g)} \rightarrow CO_{2\,(g)} \qquad \Delta H = -283\ kJ$$

$$2(CH_4 + \cancel{2}O_2 \rightarrow \cancel{CO_2} + 2H_2O) \qquad \Delta H = 2(-802\ kJ)$$

$$2(\cancel{CO_2} \rightarrow CO + \tfrac{1}{2}O_2) \qquad \Delta H = 2(283\ kJ)$$

$$2CH_4 + 3O_2 \rightarrow 2CO + 4H_2O \qquad \Delta H = \underline{-1{,}038\ kJ}$$

8. (5 pts) Without doing any calculations, predict whether ΔS is positive, negative for each of the following processes, assuming each occurs at constant temperature:

(a) $2\,K_{(s)} + F_{2\,(g)} \rightarrow 2\,KF_{(s)}$ ΔS (−)

(b) Hot air expanding ΔS (+)

(c) $2NO_{2\,(g)} \rightarrow 2\,NO_{(g)} + O_{2\,(g)}$ ΔS (+)

(d) Dew forming ΔS (−)

(e) $2\,CO_{(g)} + O_{2\,(g)} \rightarrow 2\,CO_{2\,(g)}$ ΔS (−)

9. (4 pts) When a system is at equilibrium,
(a) the reverse process is spontaneous, but the forward process is not.
(b) the forward and reverse processes are both spontaneous.
(c) the forward process is spontaneous, but the reverse process is not.
(d) the process is not spontaneous in either direction.

10. (16 pts) Consider the following reaction:

$$H_2SO_{4\,(l)} \leftrightarrow SO_{3\,(g)} + H_2O_{(l)}$$

	$H_2SO_{4\,(l)}$	$SO_{3\,(g)}$	$H_2O_{(l)}$
S° (J/mol.K)	+156.90	+256.66	+69.94
ΔH_f° (kJ/mol)	-814.0	-396.0	-285.84

(a) Calculate ΔS° for the reaction using the given S° values. **Explain the sign of ΔS°**. (6 pts)

$$\Delta S^\circ = \left[\left(1\,mol \times 256.66\,\tfrac{J}{mol\cdot K}\right) + \left(1\,mol \times 69.94\,\tfrac{J}{mol\cdot K}\right)\right] - \left(1\,mol \times 156.90\,\tfrac{J}{mol\cdot K}\right)$$

$\Delta S^\circ = 169.70\ J/K$

ΔS is positive because there are more gases in the products.

(b) Determine ΔH° for the reaction. **Explain the sign of ΔH°**. (6 pts)

$$\Delta H^\circ = \left[\left(1\,mol \times -396.0\,kJ/mol\right) + \left(1\,mol \times -285.84\,kJ/mol\right)\right] - \left(1\,mol \times -814.0\,kJ/mol\right)$$

$\Delta H^\circ = 132.2\ kJ$

ΔH is positive ⇐ endothermic process.

(c) Calculate ΔG° for the reaction. Is the reaction spontaneous, as written, at 298K? Why? (6 pts)

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= 132.16\ \text{kJ} - (298\text{K})(0.16970\ \text{kJ/K})$$

$$\Delta G^\circ = 81.6\ \text{kJ}$$

$\Delta G^\circ > 0$ Process is nonspontaneous.

(d) Determine the equilibrium constant at 298K. What reaction is favored (forward or reverse)? Why? (6 pts)

$$K = e^{-\frac{\Delta G^\circ}{RT}}$$

$$\frac{\Delta G^\circ}{RT} = \frac{81.6\ \text{kJ}}{(8.314\times10^{-3}\ \frac{\text{kJ}}{\text{mol}\cdot\text{K}})(298\text{K})}$$

$$K = e^{-32.9}$$

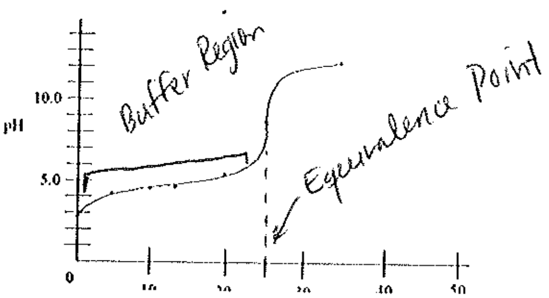
$$\frac{\Delta G^\circ}{RT} = 32.9$$

$$K = 5\times10^{-15}$$

$K < 1$ Reverse is favored

Bonus:

This figure shows the titration curve of a weak acid. Show (label) the equivalence point of the titration and the buffer region in the graph. (2pts)



Equa